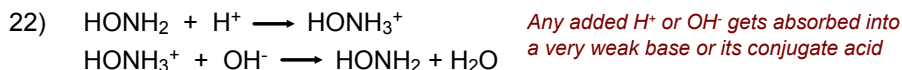


Assignment 15.1 Questions 22, 24, 26, 32, 40, 44, 46, 48



24) a. $K_b = 1.1 \times 10^{-8} = \frac{x^2}{0.100}$ b. $K_a = K_w / K_b = 9.1 \times 10^{-7}$
 $1.1 \times 10^{-9} = x^2$ $K_a = 9.1 \times 10^{-7} = \frac{x^2}{0.100}$
 $x = [\text{OH}^-] = 3.3 \times 10^{-5} \text{ M}$ $9.1 \times 10^{-8} = x^2$
 $\text{pH} = 14 - 4.48 = \mathbf{9.52}$ $x = [\text{H}^+] = 3.0 \times 10^{-4} \text{ M}$
 $\text{pH} = \mathbf{3.52}$
c. $\text{pH} = \mathbf{7.00}$
d. $\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]} = \log [9.1 \times 10^{-7}] + \log \frac{[0.100]}{[0.100]} = \mathbf{6.04}$

26) a. $\% \text{ ionization} = \frac{3.3 \times 10^{-5}}{0.100} \times 100 = \mathbf{0.033 \%}$
d. $\text{pOH} = 14 - 6.04 = 7.96$, $[\text{OH}^-] = 10^{-7.96} = 1.1 \times 10^{-8} \text{ M}$
 $\% \text{ ionization} = \frac{1.1 \times 10^{-8}}{0.100} \times 100 = \mathbf{1.1 \times 10^{-5} \%}$

*This buffer decreases the ionization of OH^- by a factor of 3000x.
Almost all of the extra OH^- reacts with the extra unreacted weak acid.
This is the common ion effect.*

32) Only trial d is a buffered solution. It has both a weak base (HONH_2) and its weak conjugate acid (HONH_3^+). If a strong acid is added, it will be mostly neutralized by the weak base. If a strong base is added, it will be mostly neutralized by the weak acid.

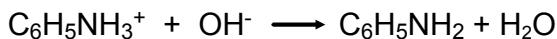
40) a. Solve for pK_b and pK_a
 $\text{pK}_b = -\log (3.8 \times 10^{-10}) = 9.42$
 $\text{pK}_a = 14 - 3.42 = 4.58$

$\text{pH} = \text{pK}_a + \log \frac{[\text{C}_6\text{H}_5\text{NH}_2]}{[\text{C}_6\text{H}_5\text{NH}_3^+]}$
 $4.20 = 4.58 + \log \frac{[0.50]}{[\text{C}_6\text{H}_5\text{NH}_3^+]}$
 $-0.38 = \log \frac{[0.50]}{[\text{C}_6\text{H}_5\text{NH}_3^+]}$

$0.417 = \frac{[0.50]}{[\text{C}_6\text{H}_5\text{NH}_3^+]}$
 $[\text{C}_6\text{H}_5\text{NH}_3^+] = \mathbf{1.2 \text{ M}}$

b. $\frac{4.0\text{g NaOH}}{1 \text{ L}} \left| \frac{1 \text{ mol}}{40.00\text{g}} \right| = 0.10 \text{ mol/L NaOH}$

The 0.10 mol of OH^- will be neutralized by 0.10 mole of $\text{C}_6\text{H}_5\text{NH}_3^+$:



	$\text{C}_6\text{H}_5\text{NH}_3^+$	OH^-	$\text{C}_6\text{H}_5\text{NH}_2$
I	1.2	0.10	0.50
C	-0.10	-0.10	+ 0.10
E	1.1	0	0.60

$\text{pH} = 4.58 + \log \frac{[0.60]}{[1.1]}$
 $\mathbf{\text{pH} = 4.32}$

44) a. $\text{pH} = \text{pK}_a + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$

$$7.40 = 6.37 + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$1.03 = \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$10.7 = \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

(inversely) $\frac{[\text{H}_2\text{CO}_3]}{[\text{HCO}_3^-]} = \mathbf{0.093}$

b. $\text{pH} = \text{pK}_a + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$

$$7.15 = 7.21 + \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

$$-0.06 = \log \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

$$0.871 = \frac{[\text{HPO}_4^{2-}]}{[\text{H}_2\text{PO}_4^-]}$$

(inversely) $\frac{[\text{H}_2\text{PO}_4^-]}{[\text{HPO}_4^{2-}]} = \mathbf{1.1}$

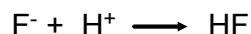
- c. A good buffer has a ratio of acid to conjugate base close to 1:1 (like H_2PO_4^- to HPO_4^{2-}). In other words, the pK_a should be close to pH. H_3PO_4 has a pK_a of 2.12, which is not close to the pH of 7.15.

- 46) A good buffer will have a pK_a close to the pH (5.00 in this case), so we want a base with a pK_b close to 9 ($\text{K}_b = 10^{-9}$). The best fit is **pyridine**.

- 48) a. No, a strong acid and its conjugate base will not work as a buffer. It would have to be a weak acid and its conjugate base.

- b. No, both are acids so they definitely won't serve as buffers.

- c. NaF is a basic salt, so it will react with the strong nitric acid:



The result is equal amounts of a weak acid (HF) and its conjugate base (F^-).

This would serve as a buffer.

	F^-	H^+	HF
I	0.4	0.2	0
C	-0.2	-0.2	+ 0.2
E	0.2	0	0.2

- d. A strong acid and a strong base will only neutralize each other (in this case there will be leftover OH^- , but they will not work as a buffer.